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Pure copper found in the Nass River Valley in British Columbia and in Alaska was pounded into plaques like the one shown here. These plaques are simply called "coppers." They were and still are important ceremonial symbols of status, wealth, and family heritage to several First Nations peoples who live along the west coast of North America.

These early uses of copper relied on native copper. However, copper occurs this way only in small amounts. It is much more common in compounds with other elements. But copper in compounds was unusable until smelting technology was developed. Smelting is the process of separating a metal from the other elements in a compound by melting. Copper was being smelted in Egypt by 4000 B.C.

Later, the discovery of tin ores caused a revolution in metallurgy. By 2500 B.C., the Sumerians, in the Middle East, had begun smelting tin ores with copper ore. They found that the metal produced from this combination was much easier to cast than pure copper because it flowed more easily and was much stronger. The copper and tin formed an alloy called bronze. An alloy is any mixture of metals. Bronze made much stronger tools and weapons than either copper or tin alone.

When people first began using metals, they thought iron was very rare. The only source seemed to be meteorites. Like copper, iron occurred more widely in mineral compounds. Eventually, a process for smelting iron became common at around 1200 B.C., and the Iron Age began. The production of iron tools revolutionized agriculture, and iron weapons did the same thing for warfare. When iron and carbon were combined in the right way, steel was produced. Steel was much stronger than iron alone. It was especially useful for hunting knives, armour, and swords.

During early times, people discovered many ways of changing and using matter but still had no basic understanding of what it was. Greater understanding was needed for the development of more advanced applications.

Aristotle's Description of Matter

What is matter? As early as 400 B.C., Greek philosophers were attempting to answer this question. They considered the idea that fundamental types of matter or elements could be combined to produce the incredible variety of substances that we see around us. The philosopher Aristotle believed that all matter was composed of combinations of fire, earth, water, and air.

Another question was whether matter could be divided into infinitely smaller and smaller pieces. Or did it become indivisible at some point? Aristotle thought fire, earth, water, and air were all continuous, which meant that there was no such thing as a smallest piece. Democritus, another Greek philosopher, had a different idea. He proposed that matter was made up of tiny particles that could not be divided into smaller pieces. He called them *atomos*, meaning "indivisible." The Greeks did not perform experiments to test their ideas, and it may not have occurred to them to try. Scientific investigation, based on experimentation, did not yet exist.

Democritus's idea of matter being made up of tiny particles was closer to what we know of the structure of matter today. However, Aristotle was better known and well respected at the time. His idea of matter being made up of fire, earth, air, and water was accepted for the next 2000 years.

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Aristotle's elements were very different from the elements that we talk about in chemistry class. Each of his elements—fire, earth, water, and air—also had two main features. Fire was dry and hot; earth was dry and cold; water was wet and cold; and air was wet and hot.

Alchemy

A scientific process of investigation did not exist, but that did not mean that no one was doing experiments. Alchemists like the one in Figure A1.13 were experimenting with matter. Alchemy was a combination of science and magic. Many alchemists hoped to get rich quickly by turning cheap metals, such as lead, into gold. Because of this, alchemists were secretive about their work. This meant that scientific knowledge was slow to develop, because people did not work together or share information about their discoveries. When an alchemist died, his knowledge disappeared with him. The same information had to be discovered again by others. Sometimes, this information was not rediscovered for hundreds of years.

Alchemy was not a real science, but it did contribute to the development of chemistry. Many important scientific advances were made during this period. Mercury was discovered, and procedures for making mineral acids, such as hydrochloric acid, were developed. Alchemists also developed or improved laboratory equipment, such as glassware and the distillation apparatus.

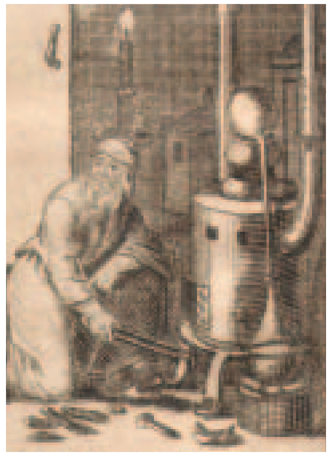

FIGURE A1.13 Alchemists developed many experimental techniques used later by chemists to make important discoveries. Chemistry developed very slowly during this early period.

In the late 1500s, people began to ask questions about how to investigate the natural world. They wondered if there was a specific procedure that could lead to a better understanding of the world. This procedure would help to ensure that results and conclusions were meaningful and true. Gradually, the scientific process that involves experimentation, observation, and forming conclusions developed. Student Reference 2: The Inquiry Process will help you use this process in your own investigations.

Like the Greek philosophers, early scientists were interested in the question: What is matter? They began to debate if the “atomos” proposed by Democritus thousands of years earlier existed. They called these tiny particles **atoms**.

Developing Hypotheses about Matter

Many scientists contributed evidence that led to our understanding of atoms. One of the first was the Irish scientist Robert Boyle, who lived from 1627 to 1691. Boyle measured relationships between volume and pressure of gases. From his experiments, he concluded that gases are made up of tiny particles that group together to make different substances. This is similar to today’s theories, but we have a much better idea of what these basic particles are.

The French scientist Antoine Lavoisier (1743–1794) measured the masses of the substances that reacted together and the substances produced in many chemical reactions. He discovered that mass is neither produced nor lost during a chemical reaction. He called this the **law of conservation of mass**. You will learn more about this law in section A3.0.

During this early period, many scientists were investigating matter, and many models of atoms were proposed. However, four classic models are always discussed because they are examples of the scientific process at its best. Each model is founded on an experiment. Although each model is attributed to one person, all the scientists applied their own insights to pre-existing ideas. The works of Dalton, Thomson, Rutherford, and Bohr illustrate the role of evidence in the development of the model of the atom.

Minds On... Atomic Models

As you read about the different models of the atom, record the development in the understanding of the atom by showing the *differences* from one model to the next. Compare Dalton's model to Aristotle's ideas, which were replaced by Dalton's work.

FIGURE A1.14 Dalton described atoms as tiny balls. The atoms of different elements were different in size and mass.

John Dalton

John Dalton (1766–1844) was an English chemist and physicist who made many contributions to chemistry. He based his model of the atom on experiments he did in combining elements.

Dalton imagined that all atoms were like small spheres, but that they could have different properties. They varied in size, mass, or colour. Figure A1.14 shows how Dalton imagined atoms to look.

Dalton used the following model to explain matter:

- All matter is made of small, indivisible particles called atoms.
- All the atoms of an element are identical in properties such as size and mass.
- Atoms of different elements have different properties.
- Atoms of different elements can combine in specific fixed ratios to form new substances.

J. J. Thomson

Joseph John Thomson (1856–1940) was an English physicist who discovered the **electron**.

In the 1890s, Thomson was experimenting with beams of particles produced in a vacuum tube like the one shown in Figure A1.15. Thomson's experiments showed that the beam was made of negative charges. By testing many different elements, he showed that they all produced the same type of beam. This suggested that atoms of different elements contained smaller particles that were identical.

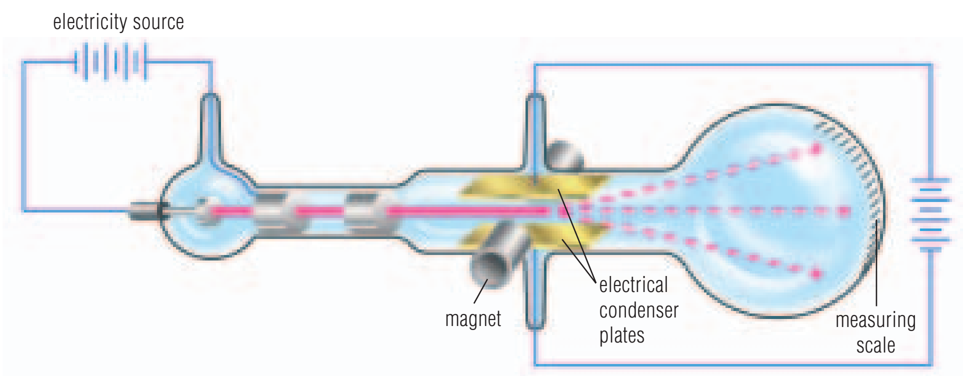

FIGURE A1.15 J. J. Thomson used this apparatus to produce beams of particles.

Thomson used his experimental evidence to develop a new model of the atom (Figure A1.16). His model stated that all atoms are made of smaller subatomic particles put together in different combinations to make the different elements. He suggested that an atom was a sphere of positive charge in which negative particles were imbedded. The negatively charged particles were called electrons. The Japanese scientist H. Nagaoka proposed a different model in 1904. He placed the electrons on the outside of the sphere. There they travelled around the central sphere in a pattern like the rings around the planet Saturn. Both models were useful explanations but neither one remained acceptable for very long.

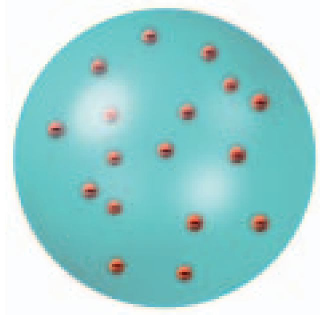

FIGURE A1.16 In Thomson's model, the atom was made up of a positively charged sphere with negative particles embedded in it.

Ernest Rutherford

Ernest Rutherford (1871–1937) began working with radioactive substances in England with J. J. Thomson. Radioactive substances release energy or charged particles. Later, Rutherford did research at McGill University in Montreal, where he performed an experiment that led to the discovery of the nucleus of the atom. He had a radioactive material encased in lead with one small opening. This material released positively charged particles which he aimed at a thin sheet of gold foil (Figure A1.17). Using Thomson's model of the atom, he predicted that all the high-speed particles would pass right through the foil. The gold atoms would either have no effect on the particles or would deflect them slightly. This is exactly what happened to most of the particles. However, a few—about 1 in 10 000—bounced back, and a few others were sharply deflected. This was entirely unexpected.

Rutherford compared this result to firing a cannon ball at tissue paper, and seeing the cannon ball bounce back occasionally! Rutherford knew that even though the unexpected results happened rarely, they still meant that Thomson's model was wrong. Rutherford developed his own model (Figure A1.18). He suggested that an atom is mainly empty space through which the positive particles could pass, but that each atom had a tiny, positively charged core. This dense core of positive charge was so strong that it was causing some of the positively charged high-speed particles to bounce back. Electrons move through the rest of the atom's volume. Rutherford called the small dense centre the **nucleus**. He calculated the size of the nucleus to be about 1/10 000 of the size of the atom. This is like the size of an ant in a football field. Rutherford received the Nobel Prize in chemistry in 1908 for his work on radioactivity.

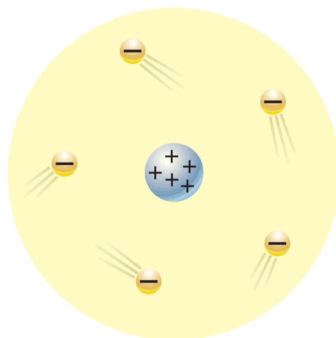

FIGURE A1.18 Rutherford's model of the atom had a tiny positively charged nucleus.

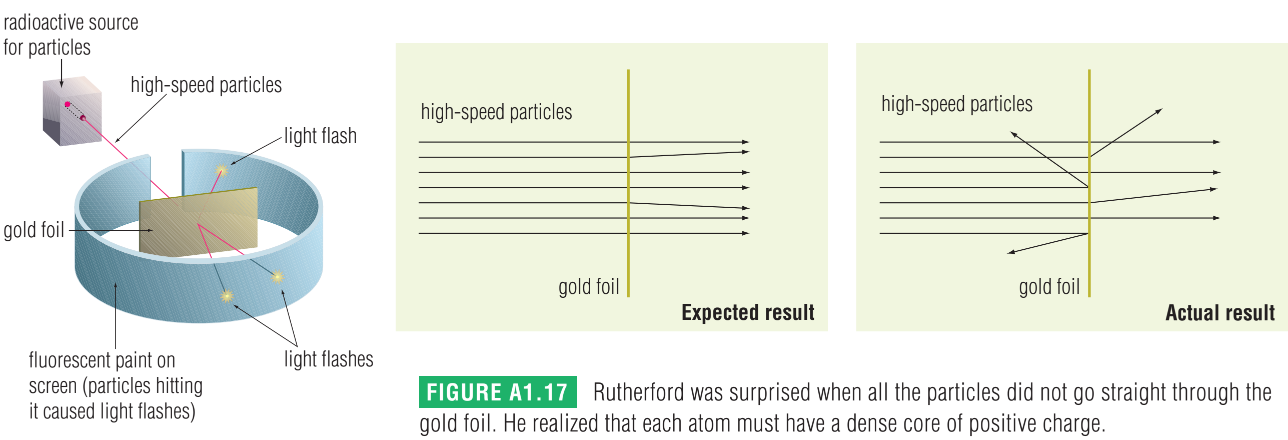

FIGURE A1.17 Rutherford was surprised when all the particles did not go straight through the gold foil. He realized that each atom must have a dense core of positive charge.

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Neon signs glow because of electrons changing energy levels. Electricity is used to raise electrons to high energy levels in the atoms of the neon gas. When the electrons move back down to lower energy levels, they release the energy, much of it in the form of visible light.

Neils Bohr

Neils Bohr (1885–1962) was a Danish physicist who worked under Rutherford in England. He proposed that electrons surrounded the nucleus in specific energy levels (Figure A1.19). He found evidence for these energy levels by examining the light released by hydrogen atoms when they are made to glow in a tube.

Figure A1.20 shows the different colours of light emitted by hydrogen atoms. The individual bands of light correspond to gaps between the energy levels of the electrons. When electrons fall from higher energy levels to lower energy levels, they release a particular colour of light. From these colours, it is possible to identify the energy levels in atoms of all the elements in the periodic table. Figure A1.21 shows the relationship between the colours of the hydrogen emission spectrum and energy levels in atoms.

Bohr's experiments also partly explained why the negatively charged electrons do not merge with the positively charged nucleus. The reason is that electrons cannot fall below the lowest energy level. Thus an electron cannot fall into a nucleus under normal circumstances.

*re*SEARCH

Rutherford knew that the nucleus is composed of many positive charges that would tend to fly apart. This implied the existence of some unknown force capable of holding them together. Investigate the strong nuclear force that holds protons and neutrons together in an atom. Begin your search at www.pearsoned.ca/school/science10

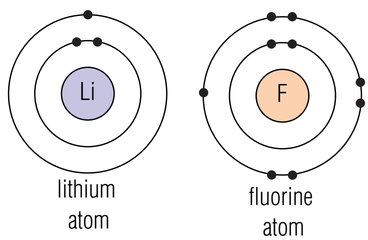


FIGURE A1.19 Bohr's model of the atom shows electrons at different energy levels orbiting the nucleus.

FIGURE A1.20 The hydrogen emission spectrum. Bohr used the range of light emitted by hydrogen atoms in his studies of the atom.

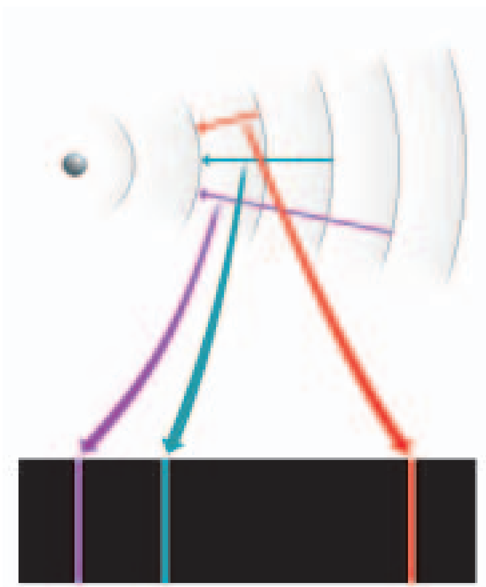


FIGURE A1.21 Electron energy levels and the hydrogen emission spectrum. When an electron falls from the third energy level to the second energy level, red light is emitted. When it falls from the fourth energy level to the second energy level, blue-green light is emitted. Similarly, a fall from the fifth to the second energy level emits violet light.

The Quantum Mechanical Model of the Atom

Today's model of the atom (Figure A1.22) is based on a theory called quantum mechanics. This abstract model is difficult to visualize. It uses mathematical probability to describe how electrons exist in atoms. Each electron can be thought of as a "cloud" of negative charge, instead of a tiny negative particle. Rather than thinking of the electron as a small particle moving quickly through a space, as Bohr originally did, the whole idea of electron movement is abandoned. Instead, electrons "occupy" the whole space all at once at different energy levels.

The electron cloud surrounds a nucleus containing two types of particles called **nucleons: protons** and **neutrons**. Protons have a positive electrical charge, and neutrons have no electrical charge. This model of the atom could change in the future as scientists learn more about the atom and subatomic particles.

FIGURE A1.22 In the model of the atom that scientists use today, electrons form a cloud around the nucleus.

A1.3 Check and Reflect

Knowledge

1. List three methods for preserving food. For each method, state if it is a physical process or a chemical process.
2. Of the five metals known to early peoples, iron was the rarest. Why?
3. How did the discovery of tin ores affect the smelting of copper?
4. List three examples of advances in chemistry made by the alchemists.
5. List the four basic ideas proposed by Dalton in describing the nature of matter.
6. What particle did J. J. Thomson discover? Where did this particle fit in his model of the atom?
7. Describe the evidence that led Rutherford to the discovery of the nucleus.
8. How did Bohr use light emitted from atoms to decide that electrons existed in specific energy levels?

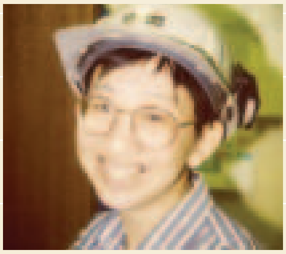
Applications

9. The First Nations people of North America used copper extensively. However, there is no evidence that they developed smelting technology. Why do you think they did not need to develop this technology?
10. Sketch four diagrams of models of the atom proposed by Dalton, Thomson, Rutherford, and Bohr. Label the important parts of each.
11. Think about the quantum mechanical model of the atom. Why is it incorrect to think of the electrons in the atom as being like planets that move around the Sun?

Extensions

12. Write a conversation between Democritus and Aristotle in which they debate their different ideas about the composition of matter.
13. Rutherford was a student in J. J. Thomson's laboratory. Write a letter from Rutherford to Thomson in which he explains his new experiments that seem to make Thomson's model of the atom obsolete.

Career and Profile



Pauline Lee is a chemical engineer at Celanese Canada Inc. in Edmonton. She was born in Hong Kong and moved to Canada with her family when she was eight. She graduated from Innisfail High School in Innisfail, Alberta.

Chemical Engineer

Chemical engineers work in careers related to the production and use of chemicals. Their jobs involve problem-solving and make use of chemistry, mathematics, physics, and practical knowledge about applying scientific principles to specific applications. For example, a chemical engineer might work in recovering and recycling materials, developing pharmaceuticals, purifying water, or refining gasoline.

How did you prepare yourself for a career in chemical engineering?
I was very good at mathematics and physics and went to the University of Alberta. I began in pharmacy, but moved into chemical engineering. I really enjoyed my classes. I found that I was very good at visualizing molecules, which helped me to understand what was going on. University teaches you how to think—how to be a detective discovering and understanding things like chemical processes. It also teaches you how to solve problems.

Where have you worked as an engineer?
I worked at Sherritt International Corp. in Fort Saskatchewan as a production process engineer for eight years. They refine nickel from nickel sulfide ores and nickel recycle. I have worked for four years at Celanese Canada Inc. in the Methanol Unit as a process engineer. Methanol is a base for products, such as formaldehyde. We make methanol from natural gas and water at high temperatures and pressures using a catalyst. A process engineer helps to optimize the process. For example, energy efficiency is important—we try to use the least amount of steam possible.

How is being an engineer different from studying engineering at school?
In school, you are given a problem and have to find a solution to it. As a process engineer, you have to find out what the problem is first. You make up your own questions, and then you answer them. Finding the right question leads you to the right answer. I really like my work.

How important are environmental concerns to you?
Environmental concerns affect the way we look at our work and design our processes. We want to preserve the planet. For example, we reuse as much of our wastes as we can. Those that can't be reused are purified as much as possible. Our objective is that nothing leaves the plant in the form of wastes. As a process engineer, I am part of the team that designs and implements these goals.

1. Why is problem-solving an important part of a chemical engineer's job?
2. If you were a chemical engineer, what kind of products or processes would you like to work on? Why?

Section Review

Knowledge

1. How many fire extinguishers are in your science class and where are they?

2. Explain why standard eye glasses are not enough eye protection in the science lab.

3. Hazard symbols specify degree of hazard using three borders of different shape and colour. What are the name, shape, and colour of each degree of hazard?

4. List four of the most important safety rules in your class.

5. Explain what is meant by the term "chemical reaction." List two characteristics common to chemical reactions.

6. What is fermentation? How is it used to preserve food?

7. What is bronze? How is it made?

8. What basic law of nature did Antoine Lavoisier discover by making careful measurements of the masses of the chemicals in his experiments?

9. How did the discovery of the electron by J. J. Thomson change Dalton's model of the atom?

10. Dalton imagined that all atoms are like small spheres. In what ways did he imagine the spheres could vary?

11. How did Bohr's model of the atom begin to explain why the negatively charged electrons don't simply fall into the positively charged nucleus?

12. In the modern theory of an atom, do the electrons move within an energy level? How is an electron thought to exist in an atom?

Applications

13. What type of hazard symbol would you expect to see used for each of the following?
- a) bleach
- b) drain cleaner
- c) lube oil
- d) pesticide
- e) solid fertilizer
- f) gasoline

14. Imagine that you have discovered a new hazardous chemical. Describe your chemical, give it some characteristics, and then write a one-page MSDS for it.

15. Classify each as a compound or a mixture:
- a) hot tea
- b) methanol, $CH_3OH_{(l)}$
- c) $NaCl_{(aq)}$
- d) a cookie
- e) sugar in water

16. A radioactive material is placed in a lead shielding block with one hole leading to the outside, as shown below. What tests would you use for the following?
- a) Determine whether any radiation is coming out of the hole.
- b) Determine whether the particles of radiation have a positive charge, a negative charge, or no electric charge at all.

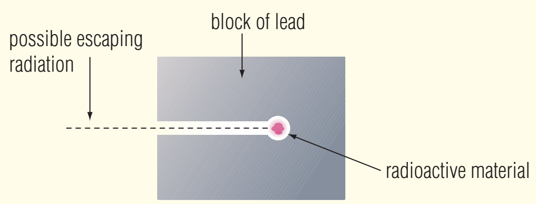


Diagram for question 16

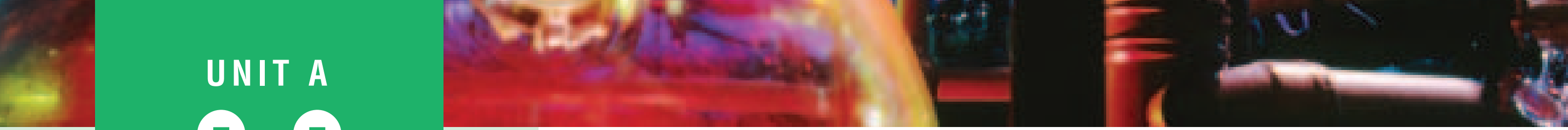
17. Rutherford concluded from his gold foil experiment that the nucleus of the atom was positively charged and most of the volume of the atom was negatively charged. The opposite conclusion would have been that the nucleus was negative and most of the volume of the atom was positive. Explain why this opposite conclusion was not made.

18. Write a letter to John Dalton explaining why some of the four points in his model of the atom are not correct.

Extensions

19. Create a model of an atom that illustrates the nucleus and the location of surrounding electrons in their energy levels.

20. Create a Web page or other electronic presentation that explains the development of the model of the atom from Dalton to Bohr.



UNIT A

2.0

Elements combine to form many substances, each with its own set of properties.

Key Concepts

In this section, you will learn about the following key concepts:

- how chemical substances meet human needs
- International Union of Pure and Applied Chemistry (IUPAC) nomenclature, ionic and molecular compounds, acids and bases

Learning Outcomes

When you have completed this section, you will be able to:

- explain, using the periodic table, how and why elements combine to form compounds in specific ratios
- explain the importance of and need for the IUPAC system of naming compounds in terms of the work that scientists do and the need to communicate clearly and precisely
- predict formulas and write names for ionic and molecular compounds and common acids
- classify ionic and molecular compounds, acids, and bases on the basis of these properties: conductivity, pH, solubility, and state
- predict whether an ionic compound is soluble in water, using a solubility chart
- relate the molecular structure of simple substances to their properties
- outline the issues related to personal and societal uses of potentially toxic or hazardous compounds
- identify examples of chemistry-based careers in your community

FIGURE A2.1 The brightly coloured fabrics we wear are the result of advances in chemistry over hundreds of years.

Think of the many different styles of clothing that people wear and how bright and varied the many colours are (Figure A2.1). There are red and yellow hues, deep blues, and dark, solid blacks. Most of the pigments used to make the dyes for these fabrics were not available as recently as 100 years ago.

At first, clothing was coloured with dyes from natural sources, such as plants. Many of these colours were not as bright or intense as the ones we see today. An understanding of the elements and how they combine led to the invention and mass production of chemicals for dyes.

In this section, you will review the elements and the periodic table. You will also review and learn more about the structure of the atoms that make up the different elements. This will lead to a study of how atoms combine chemically by gaining, losing, or sharing electrons. You will learn how to name substances and how to categorize them based on their properties. At the end of this section, you will consider how toxic and hazardous chemicals affect each of us and the environment.

A 2.1 The Periodic Table and Atomic Structure

We live in a world that contains a vast number of different materials. Yet the components of all these many materials can be separated into about 115 basic building blocks called the **elements**. Recall from earlier science studies that elements are substances that cannot be broken down into other substances. Some of the most familiar elements are carbon, oxygen, and gold.

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The blue colour in blue jeans comes from the pigment copper phthalocyanine, a synthetic pigment made from the elements copper, carbon, hydrogen, and nitrogen.

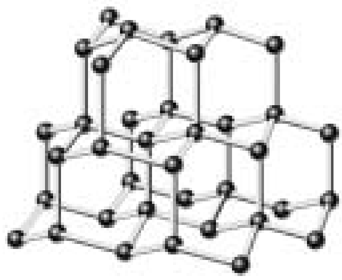
The Elements

There are about 90 naturally occurring elements, and another 25 synthetic elements. Based on their properties, all the elements can be divided into three classes: metals, non-metals, and metalloids.

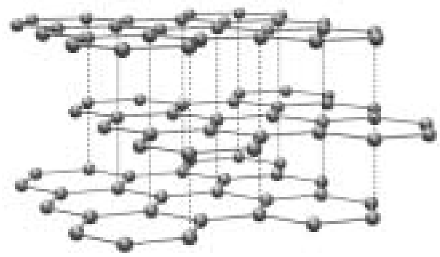
Metals

Most of the elements are **metals**. Most are silver or grey in colour and shiny. They are all good conductors of electricity and heat. They are also **malleable** and **ductile**. Malleable means that they can be beaten or rolled into sheets without crumbling. Ductile means they can be stretched into long wires.

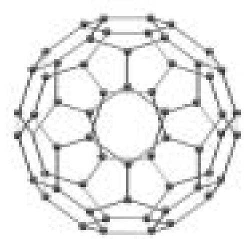
Metals have many other properties in common, although there are differences. For example, most metals are solids at room temperature (25°C). Mercury is the exception. It melts at –39°C. Another variable characteristic is how strongly metals react with other substances. Some metals, such as sodium, are highly reactive with air and water. Others such as platinum and gold are **inert**, or unreactive, except with the most corrosive acids.



(a) Diamond has a three-dimensional web of bonds, which makes it very hard.



(b) Graphite has a two-dimensional network of bonds. It is very soft.



(c) Buckminsterfullerene forms spheres. It is also soft.

FIGURE A2.2 Carbon exists in three common forms.

Non-Metals

Only 17 elements are **non-metals**. They are grouped together mainly because of their lack of resemblance to metals, rather than their similarities to each other. For example, 11 of the non-metals are gases at 25°C, 5 are solids, and 1, bromine, is a red-brown liquid. There is also tremendous variation in colour. Fluorine is pale green, and chlorine is yellow. Iodine is violet. Some non-metals exist in different forms. For example, phosphorus has a red form and a white form. Both forms are stable at room temperature. Carbon can exist in three forms as shown in Figure A2.2.

Some non-metals are highly reactive. Fluorine, for example, can etch glass. Noble gases, such as helium, are generally unreactive. About half of the non-metals exist at 25°C as connected groups of atoms called **molecules**, such as oxygen, O_2. Others, such as neon, exist only as individual atoms. You will learn more about molecules in sections A2.2 and A2.3.

Metalloids

The remaining elements are called **metalloids**, and they have properties that are intermediate between metals and non-metals. For example, some metalloids conduct electricity, but not very well. Silicon, used in the manufacture of computer chips, is a metalloid. Boron and arsenic are also metalloids.

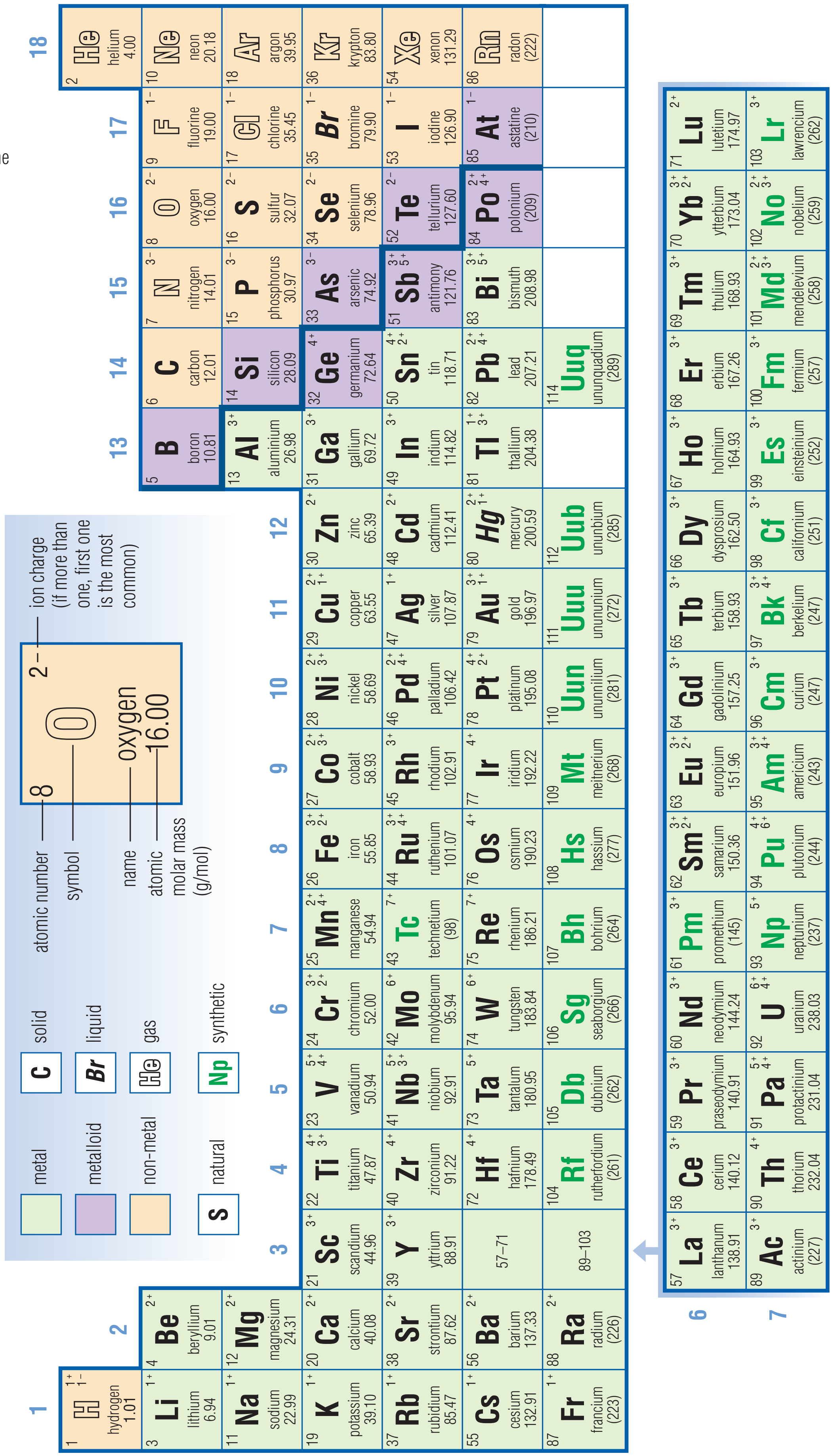


FIGURE A2.3 The periodic table displays the elements in an organized chart.

The Periodic Table

The periodic table organizes all the elements according to their chemical properties (Figure A2.3). Notice that the metals are located on the left side and centre of the table, and the non-metals are on the far right. In between are the metalloids. One exception is hydrogen. It is a non-metal, but it is located at the left side because it often behaves like a metal in chemical reactions.

Each box in the table shows the name and symbol for each element. The symbol is often an abbreviation derived from the element's name. You can see that carbon's symbol is C. Some of the metals have been known for thousands of years. Their symbols are derived from their original Latin names. For example, the symbol for lead is Pb. It is derived from the Latin word for lead, *plumbum*. The English word "plumber" is also derived from *plumbum* because Roman plumbers used lead piping.

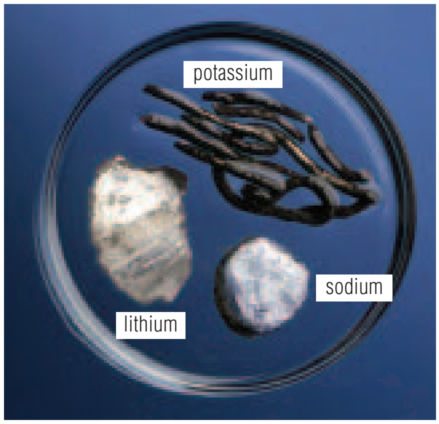


FIGURE A2.4 Lithium, sodium, and potassium are alkali metals. The alkali metals are group 1 on the periodic table.

Periods and Families

The periodic table is organized into rows and columns. Each horizontal line or row is called a **period**. The periods are numbered from 1 to 7. Hydrogen and helium make up the first period. Each vertical column forms a **group** or **family** of elements, numbered from 1 to 18.

Chemical families are groups of elements that have similar chemical and physical properties. For example, group 1 is located in the column at the far left of the table, and includes lithium, sodium, and potassium (Figure A2.4). Called the **alkali metals**, they are all soft, shiny, and silver in colour, and very reactive with water. Their compounds tend to be white solids that are soluble in water. Recall that a compound is a chemical combination of two or more elements in a specific ratio. Next to them is group 2, which includes magnesium and calcium (Figure A2.5). They are called the **alkaline-earth metals**. They are shiny and silver but are not as soft as the alkali metals. Their compounds tend to be white, but they are less soluble than compounds formed by the alkali metals.

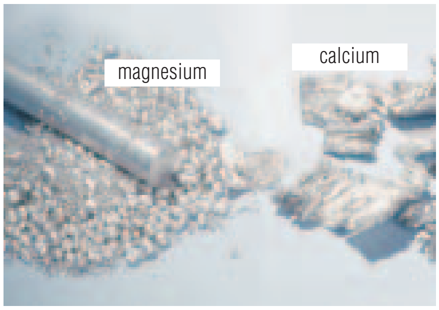


FIGURE A2.5 Magnesium and calcium are alkaline-earth metals. They are group 2 in the periodic table.

Moving to the right side of the periodic table, group 18 is a column that contains helium, neon, and argon. These are the **noble gases**. They are very unreactive. Helium has a very low density, which is why helium-filled balloons float, and its non-reactivity means it cannot catch fire. Floating party balloons are filled with helium.

The elements in group 17, just to the left of the noble gases, are called the **halogens.** This family of non-metals consists of the elements fluorine, chlorine, bromine, and iodine (Figure A2.6). These elements are poisonous and react readily with the alkali metals to form **salts**, such as sodium chloride (table salt). Salts are compounds produced in neutralization reactions between acids and bases. You will learn more about salts, acids, and bases in section A2.4.

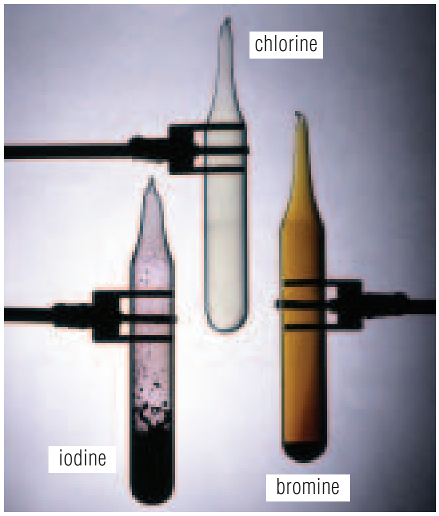


FIGURE A2.6 Iodine, chlorine, and bromine are halogens. The halogens are group 17 on the periodic table.

info **BIT**

During a supernova, a star with at least 10 times the mass of our Sun explodes. However, at the centre of the star, the explosion is *inward*, causing the core to collapse. All the atoms, most of which are iron, collapse into neutrons. The result is called a neutron star, only 20 km in diameter. The material in it has a density equivalent to collapsing a room full of solid iron into a sphere the diameter of the period at the end of this sentence.

Atomic Theory

When you write with a pencil, tiny flakes of graphite break off to form the pencil mark. Each flake is pure carbon—made up of millions of carbon atoms. The smallest possible piece of graphite is a single carbon atom. An **atom** is the smallest part of an element that still has the properties of the element. A typical atom is very small—only about 10^{-10} m in diameter. This size is hard to imagine, but an analogy can help. An orange contains many carbon atoms as well as other atoms. Imagine that all the atoms in an orange increase in size until a single atom becomes the size of the original orange. How big is the whole orange after all its atoms have expanded? If a carbon atom became the size of an orange, then the orange that contained it would become the size of the whole Earth!

Subatomic Particles

Three kinds of subatomic particles are: electrons, protons, and neutrons. Recall that electrons are negatively charged particles, and protons are positively charged particles. Neutrons are neutral particles. They have no electrical charge. All the protons and neutrons are gathered together in a tiny region at the atom's centre called the nucleus. The nucleus is so small that 10 000 nuclei (plural of nucleus) in a row would fit once across the diameter of an atom.

Despite their small size, protons and neutrons account for more than 99.9% of the total mass of an atom. This was the amazing fact that Ernest Rutherford discovered with his gold foil experiment. For example, a piece of iron metal, such as a fork, feels very solid, and its mass is detectable when you pick it up. However, more than 99.9% of the volume has virtually no mass at all. Imagine a room the size of a typical classroom filled completely with a huge block of iron. Then imagine taking all the nuclei out of all the iron atoms and placing them side by side, and touching. How big would the volume of all the nuclei be? It would be about the size of the period at the end of this sentence. Yet its mass would be almost equal to the mass of the room full of iron!

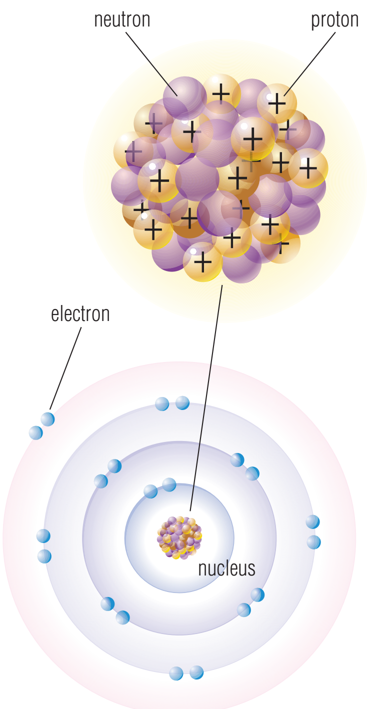


FIGURE A2.7 Electrons occupy most of the volume of an atom. Most of the mass is in the nucleus. A calcium atom (shown here) has 20 electrons, 20 protons, and 20 neutrons.

Energy Levels

Electrons take up most of the volume of an atom (over 99.9%), and they occupy specific energy levels. An **energy level** can be thought of as a region of space near a nucleus that may be empty or may contain electrons. Electrons in energy levels nearest the nucleus have the lowest energy. Electrons in energy levels farther away from the nucleus have more energy. Electrons in the lowest energy levels are the most tightly held in the atom because they are closest to the positively charged nucleus.

The number of electrons that can exist in the different energy levels varies. The lowest energy level is the one closest to the nucleus. It can hold only 2 electrons. The next energy level is larger and farther from the nucleus. It can hold up to 8 electrons. Think of these energy levels as being like spheres that add new layers to the outside of the atom, just as an onion has layers. The third energy level also has room for up to 8 electrons. It is common to discuss the electron arrangement in atoms up to 20 electrons, which is calcium. Beyond that, the pattern is more complicated. For calcium, the electron distribution is 2, 8, 8, 2, as shown in Figure A2.7.

An energy level can be empty, partly filled, or completely filled. Partly filled energy levels from two different atoms can overlap, and a pair of electrons can exist in both of them at once. This is the basis for chemical bonding, which you will study later in this section.

Electrons and protons are attracted to each other because they have opposite charges. However, they cannot completely come together under normal circumstances. This is fortunate, because it means that atoms cannot collapse. Some of the properties of these particles are summarized in Table A2.1.

info **BIT**

The first elements are thought to have formed during the "Big Bang" when the universe formed.

TABLE A2.1 Properties of Protons, Neutrons, and Electrons

Particle	Symbol	Charge	Mass	Location
proton	p^+	1+	1.7×10^{-24} g	nucleus
neutron	n^0	0	1.7×10^{-24} g	nucleus
electron	e^-	1−	9.1×10^{-28} g	surrounding the nucleus

Atomic Number

All atoms of an element have the same number of protons. For example, all hydrogen atoms have exactly one proton. Atoms of helium, the next element in the periodic table, have two protons—never more, never less. The **atomic number** of an element indicates the number of protons it has. This number can be used to specify an element.

Look at the periodic table in Figure A2.3 on page 30. Notice that the elements in a period are arranged according to increasing atomic number. The element with atomic number 3 is lithium. Therefore, all atoms of lithium have three protons. Beryllium has atomic number 4. All atoms of beryllium have four protons. To the right of beryllium is boron with five protons. As you move from left to right in a period, each element has one more proton in its atom.

Mass Number and Atomic Molar Mass

Atoms of the same element that contain different numbers of neutrons are called **isotopes**. For example, the most common form of hydrogen has one proton and no neutrons at all. However, about 1 in 10 000 hydrogen atoms contains one proton and one neutron. This isotope of hydrogen is called deuterium. It is sometimes also called "heavy hydrogen" because the neutron increases the mass of the atom. Deuterium is used in the production of heavy water for Canadian nuclear reactors. Hydrogen has a third isotope, called tritium, that has 1 proton and 2 neutrons.

To help distinguish between the isotopes of an element, each isotope is given a number called the **mass number**. The mass number is an integer equal to the total number of protons and neutrons in the nucleus of an atom. Electrons are not included in the mass number because their mass is so small. Oxygen has three naturally occurring isotopes. They all have the same number of protons, so they have the same atomic number (8). The most common isotope has a mass number of 16: it has 8 protons and 8 neutrons. The other two isotopes have mass numbers of 17 (8 protons and 9 neutrons) and 18 (8 protons and 10 neutrons).

The atomic symbol for an element is sometimes shown with both the mass number and atomic number as follows:

$${}^{\text{mass number}}_{\text{atomic number}}\text{element symbol}$$

Using this format, the symbols for the oxygen isotopes are:

$${}^{16}_{8}O \qquad {}^{17}_{8}O \qquad {}^{18}_{8}O$$

The isotopes and their symbols are shown in Figure A2.8.

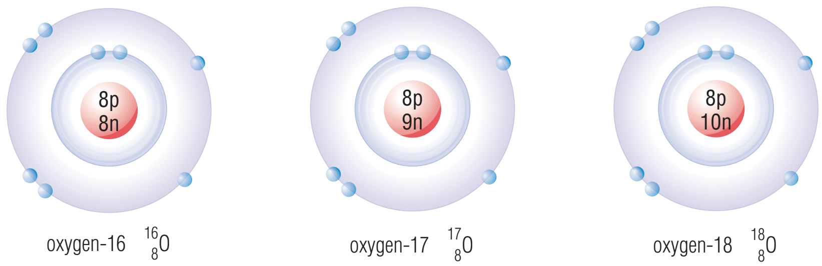


FIGURE A2.8 The three naturally occurring isotopes of oxygen. The only difference among them is the number of neutrons in their nuclei.

Note that the atomic symbol written this way does *not* give either the number of neutrons or the number of electrons directly. You can determine the number of neutrons in an atom by subtracting the atomic number from the mass number. For example, the oxygen isotope ${}^{18}_{8}O$ has a mass number of 18 and an atomic number of 8. The number of neutrons can be determined by subtracting 8 from 18.

mass number (18) − atomic number (8) = number of neutrons (10)

The **atomic molar mass** shown in the periodic table is related to the mass number. The atomic molar mass is the average mass of the element's isotopes. Isotopes of an element do not have exactly the same mass: some have slightly greater masses than others. This is because each isotope has a different number of neutrons in its nucleus.

In general, atoms are neutral, so the number of electrons in an atom equals the number of protons.

Formation of Ions

Under some circumstances, the atoms of most elements will either gain or lose one or more of their outermost electrons. The process of gaining or losing electrons is called **ionization**, and it results in the formation of an **ion**. An ion is an electrically charged atom or group of atoms. Ionization results in metals and non-metals forming compounds.

Positively charged ions are called **cations**. Most cations form when metal atoms lose electrons. When a cation is forming, the lost electrons usually move to another atom. Electrons are negatively charged. When they leave an atom, the ion that remains is positively charged because it now has more protons than electrons.

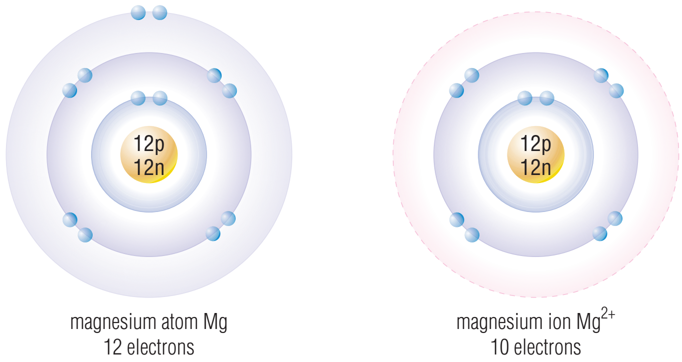


FIGURE A2.9 A magnesium atom, Mg, can lose two electrons to form a magnesium ion with a positive charge of 2+: Mg^{2+}.

For example, the periodic table shows that a sodium atom has 11 protons. Therefore, it also has 11 electrons. A sodium ion forms when the sodium atom loses one electron. This means that the positive charge in the sodium becomes one greater than the negative charge. So the sodium ion has a positive charge. An ion's charge is sometimes shown as a superscript after the atomic symbol. In this example, the sodium ion would be written as Na^+. The magnesium ion in Figure A2.9 has a charge of 2+, so it is written as Mg^{2+}.

Negatively charged ions are called **anions**. Most anions form when non-metal atoms gain electrons. For example, oxygen atoms can gain two electrons. The periodic table shows that an oxygen atom has eight protons, and therefore, it has eight electrons. By gaining two electrons, the oxygen atom becomes a negatively charged ion, called an oxide ion. It can be written as O^{2-}. Note that the name of an element's anion is written by using the first part of the element's name and changing the last part to "-ide." So the nitrogen anion N^{3-}, for example, is called nitride. The fluorine atom in Figure A2.10 forms a fluoride ion, F^-.

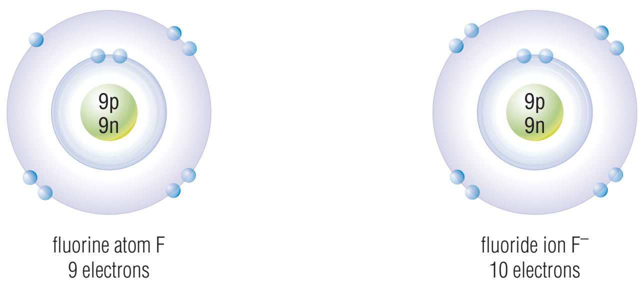


FIGURE A2.10 A fluorine atom, F, can gain one electron to form a fluoride ion, F^-.

Most ion formation takes place when metals and non-metals form ionic compounds. Metal atoms tend to form cations by losing electrons to non-metals, which form anions. You may be wondering why metals tend to lose electrons, while non-metals tend to gain them. Atoms gain or lose electrons so that they have the same number of electrons as the nearest noble gas. This makes them more stable. In chemistry, becoming more stable means becoming less reactive. For example, a sodium ion is less reactive than a sodium atom. Sodium metal contains only atoms. When it is placed in water, a vigorous reaction occurs (Figure A2.11). Table salt contains sodium ions (and chloride ions). When it is placed in water, it dissolves quietly.

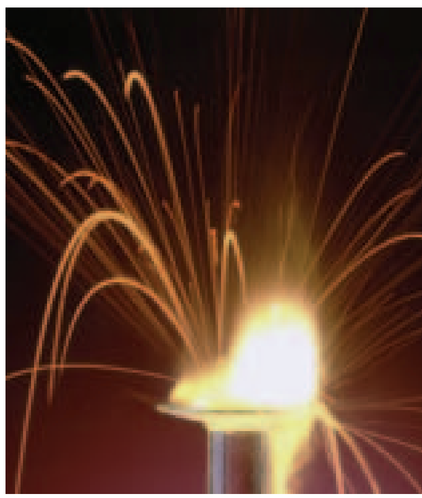


FIGURE A2.11 Sodium and water react vigorously. The gas produced ignites in a spectacular manner.

Elements Combine to Form Compounds

Recall that elements in the same group or family on the periodic table have similar physical and chemical properties. One of these properties is reactivity. An element's reactivity is related to the number of electrons in its outer energy level. Elements are most stable, or unreactive, when they have filled outer energy levels.

Recall that the noble gases (group 18 on the periodic table) are very stable. They have filled outer energy levels, as shown in Figure A2.12. The noble gases neither gain nor lose electrons. Other elements are not as stable as the noble gases. These elements gain or lose electrons. They become more stable when they have the same number of electrons in their outer energy level as the nearest noble gas does.

The electrons in the outer energy level are called **valence electrons**. The tendency to gain or lose electrons is sometimes called **valence**. The term **valence number** is commonly used to describe the number of electrons an element can gain or lose to combine with other elements. Figure A2.12 shows the number of electrons in each atom in part of the periodic table. It illustrates several striking patterns.

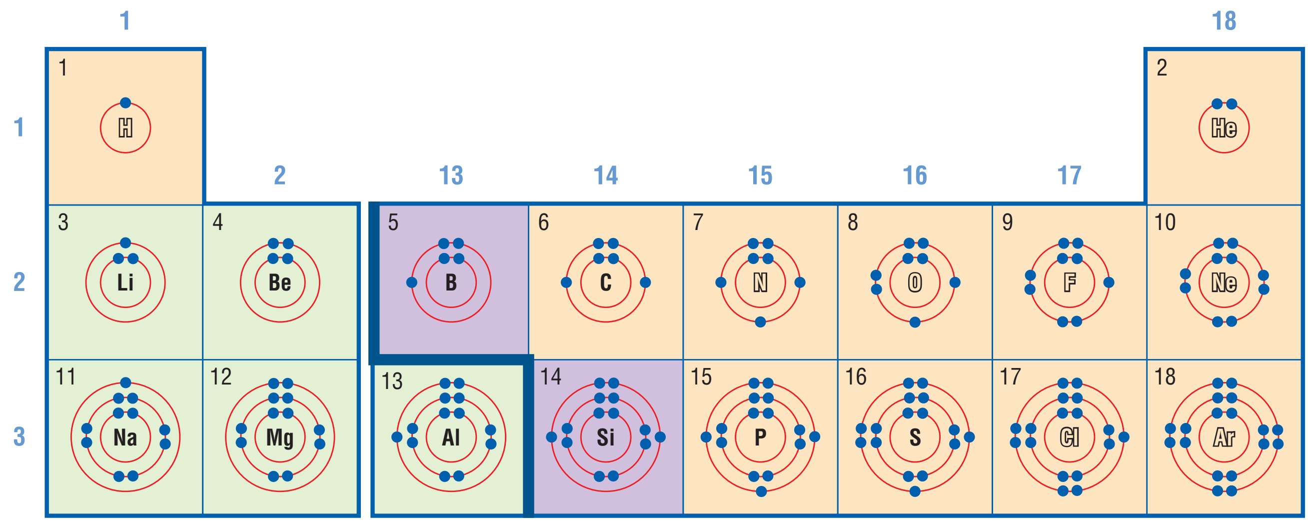


FIGURE A2.12 Part of the periodic table showing electron arrangements for the atoms of each element

Elements in the same family have the same number of valence electrons. This results in similar chemical properties. For example, the alkali metals lithium, sodium, and potassium are all soft, shiny metals that react vigorously with water. Each has one valence electron and is one atomic number away from a noble gas. Each loses one electron to form an ion with a 1+ charge. The loss of one electron gives them a filled outer energy level, just like the nearest noble gas.

Now look at the periods in Figure A2.12. Hydrogen and helium make up the first period, and lithium is the first element in the second period. From left to right across a period, atoms gain one valence electron (and one proton) with each new element. Within a period, electrons are always added to the same energy level. An element in the second period has electrons in two energy levels, and an element in the third period has electrons in three energy levels. The period number indicates the number of occupied energy levels.

Classifying Unknown Liquids

Required Skills
- Initiating and Planning
- Performing and Recording
- Analyzing and Interpreting
- Communication and Teamwork

The periodic table was founded on examining chemical and physical characteristics of materials, and arranging them according to similarities and differences. Often, the identity of the material was unknown and tests were done to see whether it matched previously known materials. If it was a pure substance (it could not be broken down further) and it had novel properties, a new element had been discovered! A similar process will be used in this activity.

In this activity, you will design and test a procedure to distinguish between five similar liquids. By carefully observing their physical and chemical properties, you can detect differences between them.

The chemicals you will use to help you distinguish between the liquids are called test **reagents** in the Materials and Equipment list below. A reagent is a substance used for identifying, measuring, or producing other substances.

The Problem

How can you distinguish between five similar liquids?

Materials and Equipment

5 liquids labelled 1 to 5

spot plate or test tubes and racks

test reagents:
- baking soda
- sodium chloride
- thymol blue solution
- calcium supplement
- tincture of iodine
- 1-cm strips of magnesium

CAUTION: Some of the substances you will be using are mildly irritating to skin; some are corrosive; and some are flammable. Note that iodine will stain skin and clothing. Handle all substances carefully.

Conduct Your Investigation

1. Make a table in your notebook like the one below to summarize your six tests of the five liquids. Once all the tests are done, you will use the table to analyze your results and write a test procedure.
2. Using a spot plate with many test wells or a set of test tubes in a test tube rack, combine each liquid with each test reagent and record your results.
3. Select the smallest set of test reagents that can reliably distinguish between all five unknown liquids.
4. Develop a test procedure. It must be a series of steps that will allow someone to take an unlabelled sample of one of the five liquids and determine which one it is. Make sure that your procedure describes the tests to be performed and how to interpret the results.
5. Have your teacher approve your procedure.

Test and Evaluate

6. Test the effectiveness of your analytical procedure with a sample of one of the unknown liquids. Use your procedure to determine which of the five liquids it is. Test more unknown liquids as time permits.
7. Follow your teacher's instructions for disposing of all the substances you have used.
8. Review your procedure based on your testing of the unknown samples. Did you need to revise it to deal with unexpected observations?

Communicate

1. Write a summary report that contains the procedure and the results of your tests on the unknown samples.

	baking soda	sodium chloride	thymol blue	calcium supplement	tincture of iodine	magnesium strip
Liquid 1						
Liquid 2						

The Octet Rule

Another way to understand the patterns by which atoms gain or lose electrons is to look at their energy levels. All the noble gases have filled energy levels. So it is possible to restate the pattern for ion formation in a new way: Atoms tend to gain or lose electrons so that they end up with completely filled energy levels. Notice in Figure A2.12 that the noble gases neon and argon each have eight electrons in their valence energy levels. The **octet rule** (also called the rule of eight) states that atoms bond in such a way as to have eight electrons in their valence energy level. ("Oct-" means eight, so an octet is a group of eight.) This is just another way to say that atoms tend to be stable with full outer energy levels. However, it is a handy rule for figuring out an element's charge or valence by looking for the octet.

For example, chlorine is just below fluorine in group 17. All atoms in this family have seven valence electrons. By the octet rule, fluorine will have an octet of electrons (eight electrons) in its valence energy level if it gains an electron to form the F^- ion (Figure A2.13). This means that a chlorine atom will also gain one electron, forming the Cl^- ion.

The exceptions to the octet rule are hydrogen, lithium, and beryllium. They each need only two electrons in their valence energy levels because their nearest noble gas, helium, has two electrons.

The situation is more complicated with transition metals. In the periodic table, these are all the metals from scandium to zinc inclusive, and any metals directly below them. All metals tend to lose electrons to become more stable, but it is difficult for atoms to lose more than about three electrons. This is because, every time an electron is lost, the remaining electrons are held more tightly by the nucleus. Gold, for example, can lose at most three electrons, to form Au^{3+}. Depending on the chemical conditions, iron can lose either two or three electrons, to form Fe^{2+} or Fe^{3+}. The elements boron, silicon, and carbon rarely form ions. Predicting the number of electrons transition metals will lose is difficult. Consult the periodic table for the ion charge for these elements. The first charge given is the most common.

reSEARCH

Since 1950, about 20 new elements have been created or discovered. Find out how new elements are created. Begin your search at www.pearsoned.ca/school/science10

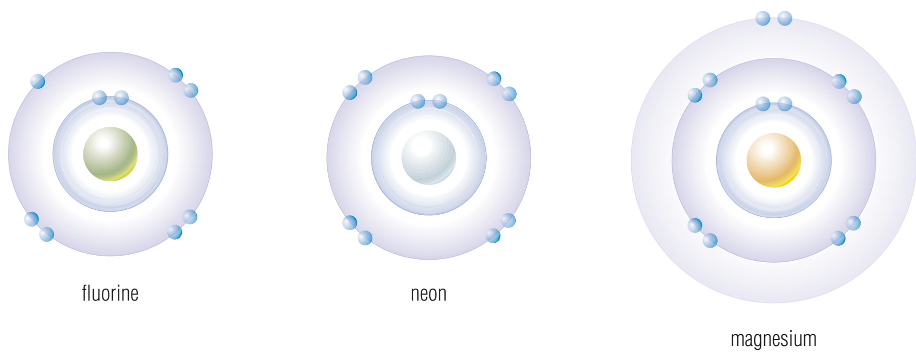


FIGURE A2.13 Fluorine has seven valence electrons; neon has eight valence electrons; and magnesium has two valence electrons. Using the octet rule, we can determine that fluorine gains one more electron; neon does not gain or lose any; and magnesium loses two to join other elements.

A2.1 Check and Reflect

Knowledge

1. List the names of four families in the periodic table. Name three elements in each family.
2. Which subatomic particle determines what the element is?
3. a) What is a valence energy level?
   b) What are valence electrons?
4. What are isotopes?
5. What is an ion? What are the special names for positive and negative ions?
6. How is the number of electrons gained or lost by atoms related to the noble gas nearest them in the periodic table?
7. State the octet rule.

Applications

8. Draw an atom that has six protons, seven neutrons, and six electrons. Using the periodic table, identify the element. Label the nucleus, the subatomic particles, and the valence energy level.
9. Two isotopes of nitrogen are nitrogen-14 and nitrogen-15. Explain how these atoms are similar and how they are different by describing their atomic structure.
10. Refer to the periodic table in Figure A2.3 on page 30 to answer the question.

    Using the octet rule, decide if each of the following elements will gain or lose electrons to become ions. State how many electrons will be involved in each case:
    a) phosphorus
    b) sodium
    c) chlorine
    d) magnesium
    e) iodine
11. Copy and complete the following table in your notebook. Use the periodic table in Figure A2.3 on page 30, as needed, to fill in the number of protons or the name of an element.

Element Name	Mass Number	Number of Protons	Number of Neutrons
calcium	41		
uranium	238		
aluminium			14
	9		5
	19	10	
iron			27

12. Copy and complete the following table in your notebook. Refer to the periodic table in Figure A2.3 on page 30.

Atom or Ion Name	Overall Charge	Number of Protons	Number of Electrons	Symbol	Number of Electrons Lost or Gained
oxygen atom				O	
oxide ion			10	O^{2-}	
potassium ion		19			
		12	10		
				F^{-}	
	2+	20			
	3+		10		

Extension

13. Using the periodic table on page 30, draw a periodic table from atomic number 1 to atomic number 20. For each element, write its symbol and the number of valence electrons in its atom. In a different colour, write down the value of the electric charge (e.g., 1^{+}, 3^{-}) on the ions of that element. Explain the pattern of ion charges either by using the octet rule or by referring to filled valence energy levels.

FIGURE A2.14 Both these materials are commonly known as chalk. Their scientific names are different and show that they are different substances.

A 2.2 Naming Ionic and Molecular Compounds

Does your classroom have chalk? You have probably used blackboard chalk. You may also have used another product that contains chalk—the antacid called TUMS. The chalk in TUMS is mixed with sweeteners and flavours so that it can be eaten easily. However, do not eat blackboard chalk for your upset stomach because blackboard chalk and the chalk in TUMS are not the same chemical at all. Blackboard chalk is mainly calcium phosphate, while the antacid chalk is calcium carbonate (Figure A2.14). This is just one example of the confusion that can result from inaccurately naming compounds. It also shows the importance of using names that provide information about the chemical composition of a substance. The term "chalk" gives no hint as to what elements are present in either compound.

The International Union of Pure and Applied Chemistry (IUPAC) is the body responsible for naming compounds. It ensures the use of a consistent, practical way of naming compounds that allows scientists to communicate clearly and precisely. Throughout this unit, you will be using IUPAC names for different types of compounds. First, we will look at ionic compounds.

info **BIT**

We throw salt on the roads in winter. But you don't throw road salt on your food! Road salt is calcium chloride. Table salt is sodium chloride. The term "salt" is not really the name of a compound. It is the name of a class of related compounds.

Ionic Compounds

Table salt—sodium chloride—is one of the most common compounds on Earth. The oceans are salty because of sodium chloride and other salts. Our cells control the amount of water they contain by controlling the concentration of salts, including sodium chloride. A high salt content in our cells draws water into them. Drawing water into the cells in our tissues removes water from our blood, causing the concentration of salts in our blood to increase. A high salt content in our blood triggers a feeling of thirst. We then have the desire to drink, and when we do, we replace the fluid that was originally drawn into the cells of our tissues. Drinking to quench our thirst adds fluids to reduce the salt concentration in our blood.

Salt is also used for electrical messaging in our bodies. Our nervous system, including our brain, is the most complex wiring system known. Salt conducts the electrical signals in our nerves because it forms ions. Recall from section A2.1 that ions are electrically charged atoms or groups of atoms. Salt belongs to a class of substances called ionic compounds.

Ionic compounds form when electrons transfer from one atom to another. For sodium chloride, positively charged sodium ions are attracted to negatively charged chloride ions. The two kinds of ions group together in an organized array called a **crystal lattice**. The lattice is made up of one sodium ion for every one chloride ion (Figure A2.15). Such a neutral unit is called a **formula unit**.

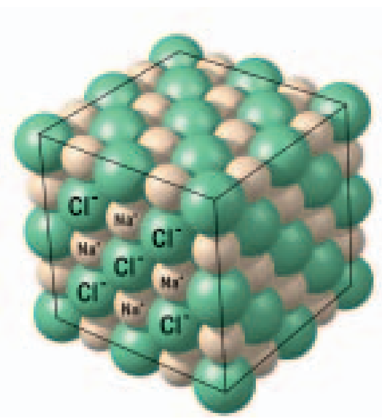


FIGURE A2.15 Table salt (NaCl) is an ionic compound. As a solid, it forms a crystal lattice.

Recall that an atom of sodium has one valence electron, and a chlorine atom has seven valence electrons. When the two elements combine, the sodium atom transfers an electron to the chlorine atom (Figure A2.16). As a result, both atoms now have full outer energy levels. Remember that the most

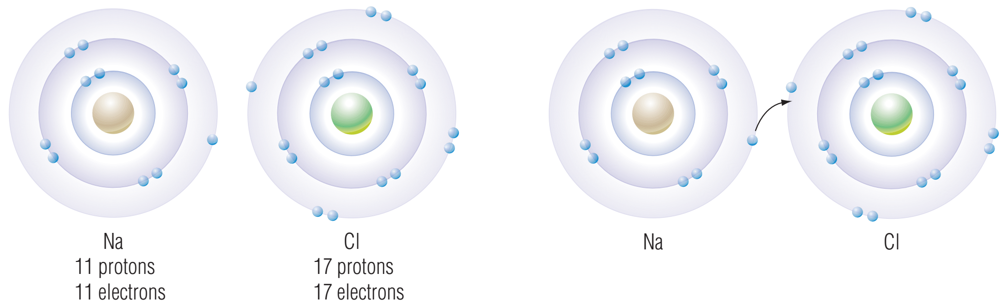


FIGURE A2.16 (a) Sodium has an atomic number of 11, so it has 1 valence electron. Chlorine has an atomic number of 17, so it has 7 valence electrons.

(b) One electron transfers from the sodium atom to the chlorine atom.

stable atoms have full outer energy levels. When sodium transfers an electron to chlorine, sodium's outer energy level is now full. When chlorine receives the electron, its outer energy level is also now full. Both elements are now stable as ions. This type of bonding is called **ionic bonding**. Ionic bonds form between atoms of metals and non-metals (Figure A2.17).

Ionic compounds have many common properties. For example, all of them are solids at room temperature. Table salt can be heated to a very high temperature without decomposing or burning. Its melting point is 800°C, which can be reached only by using a blow torch or a special oven. Ionic compounds also tend to dissolve in water, although some dissolve much better than others. Solutions of ionic compounds always conduct electricity. You will learn more about the properties of ionic compounds in section A2.3.

There are many thousands of different ionic compounds. Some have common names, such as table salt. They also have chemical names that reveal something about the elements in them. Table A2.2 gives examples of the chemical formulas and names of some common ionic compounds. Recall that the state of an element or compound is indicated by a subscript: *(s)* for solid, *(l)* for liquid, and *(g)* for gas. The subscript *(aq)* stands for aqueous. This means that the element or compound is dissolved in water.

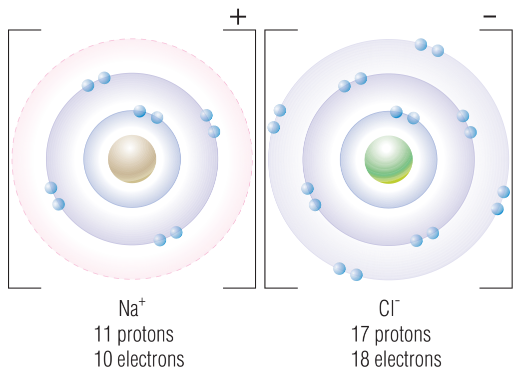


FIGURE A2.17 Sodium and chlorine form an ionic bond when their ions join to become the compound NaCl.

TABLE A2.2 Names, Formulas, and Uses of Some Common Ionic Compounds

Common Name	Formula	Chemical Name	Application
lye	$NaOH_{(s)}$	sodium hydroxide	unclogs drains
baking soda	$NaHCO_{3(s)}$	sodium hydrogencarbonate	raises breads and cakes by giving off CO_2 when heated
milk of magnesia	$Mg(OH)_{2(s)}$	magnesium hydroxide	works as antacid and laxative
table salt	$NaCl_{(s)}$	sodium chloride	adds salty taste in foods
cream of tartar	$KHC_4H_4O_{6(s)}$	potassium hydrogentartrate	mixes with baking soda to make baked goods rise

Naming Ionic Compounds

The IUPAC system of naming ionic compounds is very simple. All names of ionic compounds have two parts, because all ionic compounds are made from two parts. Every ionic compound is made up of a cation (positive) and an anion (negative). The naming rules work like this:

1. Name the cation first by using the element's name. (It is usually a metal ion.)
2. Name the anion second by using the first part of the element's name and changing the last part to "-ide." (The anion is usually a non-metal ion.)

Table A2.3 gives some examples of how ionic compounds are named.

TABLE A2.3 Examples of How Ionic Compounds Are Named

Formula	Cation	Anion	Name
$NaCl_{(s)}$	Na^+	Cl^-	sodium chloride
$BaF_{2(s)}$	Ba^{2+}	F^-	barium fluoride
$K_3N_{(s)}$	K^+	N^{3-}	potassium nitride

Formulas for Ionic Compounds

The formula of an ionic compound contains element symbols that identify each type of ion present. In some formulas, the symbols are followed by subscript numbers that indicate the ratio of ions in the compound. For example, in Table A2.3, the formula for $BaF_{2(s)}$ contains subscripts. In $BaF_{2(s)}$, there is one barium ion for every two fluoride ions. This represents the formula unit. The formula unit is the smallest amount of a substance with the composition shown by the chemical formula. It consists of positive and negative ions in the smallest whole-number ratio that results in a neutral unit in the crystal lattice of a compound. If there are no subscripts, assume that the compound has one of each ion, so the ratio is one to one in the formula unit (e.g., $NaCl_{(s)}$).

All ionic compounds are composed of an equal number of positive and negative charges. This means that the total charge of the cations must equal the total charge of the anions. In some compounds, the cation has a charge of 1+, and the anion has a charge of 1−. Recall that this is because there has been a transfer of electrons between the atoms. Consider sodium chloride. The sodium ion is Na^+ and the chloride ion is Cl^-. So only one ion of each element is needed to make the positive and negative charges equal. The ratio of sodium ions to chloride ions is one sodium ion to one chloride ion in a formula unit. Therefore, the formula is NaCl. No subscripts are needed because both ions have a charge of one.

Now consider the example of aluminium chloride, $AlCl_3$. All the charges in the formula unit must be equal. So each aluminium atom loses three electrons, and each chlorine atom gains one. The aluminium ion has a charge of 3+. The chloride ion has a charge of 1–. Therefore, every one aluminium ion combines with three chloride ions. The ratio is one cation (Al^{3+}) to three anions (Cl^-), so the formula is $AlCl_3$. Table A2.4 summarizes the two examples we have just worked through.

TABLE A2.4 Steps for Writing Formulas for Ionic Compounds

Steps	Examples	
	sodium chloride	**aluminium chloride**
1. Identify the ions and their charges.	sodium: Na^+ chloride: Cl^-	aluminium: Al^{3+} chloride: Cl^-
2. Determine the total charges needed to balance.	Na^+ : 1 Cl^- : 1	Al^{3+} : 3 Cl^- : 1+1+1 = 3
3. Note the ratio of cations to anions.	1 to 1	1 to 3
4. Use subscripts to write the formula, if needed.	NaCl	$AlCl_3$

The Method of Lowest Common Multiple

The method of lowest common multiple is another way of determining the correct formula for an ionic compound. First, find the lowest common multiple of the charges for the two ions. Then divide by the combining capacity of one ion to get the correct subscript for that ion. Repeat the process for the other ion. This ensures that the number of positive charges equals the number of negative charges, so the formula unit is electrically neutral.

Example Problem A2.1 shows the two ways of working out a formula. With the method of lowest common multiple, find the smallest number that both ion charges divide into evenly. For calcium nitride, calcium has a charge of 2+ and nitride has a charge of 3–. The lowest common multiple of 2 and 3 is 6. To find the subscript for calcium, divide 6 by 2 to get 3. For nitride, divide 6 by 3 to get 2. The formula is $Ca_3N_{2(s)}$.

Example Problem A2.1

What is the formula for calcium nitride?

Calculating the ratio:

1. Identify the ions and their charges. calcium: Ca^{2+} nitride: N^{3-}

2. Determine the total charges needed to balance. Ca^{2+}: 2 + 2 + 2 = 6; N^{3-}: 3 + 3 = 6

3. Note the ratio of cations to anions. 3 to 2

4. Use subscripts to write the formula. $Ca_3N_{2(s)}$

OR

Using the method of lowest common multiple:

1. Identify the ions and their charges: Ca is 2+, and N is 3–.

2. Find the smallest number that both charges will divide into. For Ca and N it is 6.

3. Divide each charge into the lowest common multiple, and write the numbers as subscripts: Ca: 6 ÷ 2 = 3 N: 6 ÷ 3 = 2
The formula is $Ca_3N_{2(s)}$.

Practice Problem

1. Write formulas for the following ionic compounds:

a) magnesium chloride
b) sodium sulfide
c) calcium phosphide
d) potassium nitride
e) calcium fluoride

Compounds with Multivalent Elements

Some metals have more than one stable ion. For example, iron has two stable ions: Fe^{2+} and Fe^{3+}. Elements with more than one stable ion are called **multivalent** elements. Ionic compounds containing multivalent elements must have Roman numerals in their names to indicate which ion is forming that compound. The Roman numeral is written in brackets after the element to indicate the charge. For example, chromium is multivalent, so chromium(III) sulfide indicates that the Cr^{3+} ion forms that compound. Roman numerals are not used in formulas, because you can figure out the charge on the ion by looking at the formula.

You can find the Roman numeral to use in the name of a multivalent ion by using the subscripts in the formula. For example, in $FeBr_2$, the subscript 2 after the Br is a guide to the iron ion's charge. Recall that the positive and negative charges in an ionic compound must be equal. According to this rule, only an Fe^{2+} could pair up with two Br^- to give this formula unit. $FeBr_2$ would be written out as iron(II) bromide. In $FeBr_3$, only an Fe^{3+} could pair up with three Br^- to give this formula unit. $FeBr_3$ would be written out as iron(III) bromide.

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An older system of naming multivalent ions is still commonly used that involves changing the ending of the cation. The suffix "-ous" indicates a lower charge. The suffix "-ic" indicates a higher charge. In this system, iron(II) chloride is called ferrous chloride and iron(III) chloride is called ferric chloride. The prefix "ferr" comes from the Latin word for iron, *ferrum*.

Example Problem A2.2

Write the name of the compound that has the formula $Cu_3N_{(s)}$.

1. Identify the ions that form the compound.
   $Cu^?$ copper ion N^{3-} nitride ion
2. Use the charge of the nitride ion (3−) and the rule that the total positive and negative charges in the formula unit must be equal. Three copper ions are present in the formula unit so each must have a charge of 1+.
3. Write the name of the compound.
   The name of the compound is copper(I) nitride.

Practice Problem

2. Write out the name of the following compounds:
   a) $FeCl_{3(s)}$ d) $CuF_{2(s)}$
   b) $PbO_{2(s)}$ e) $Cr_2S_{3(s)}$
   c) $Ni_2S_{3(s)}$

When writing the names of ionic compounds, either recall the charges of anions from memory or use a reference table (Table A in Student Reference 12). Use the anion's charge to find the cation's charge when the cation is multivalent. Remember that the Roman numeral is needed only if the metal element is multivalent. You can use a reference table or periodic table to find out which elements are multivalent. Select the first one listed in the periodic table if you are not given any other information. This is the most common ion for each element.

TABLE A2.5 Some Common Polyatomic Ions

Polyatomic Ion Name	Formula
ammonium	NH_4^+
carbonate	CO_3^{2-}
dihydrogenphosphate	$H_2PO_4^-$
hydrogencarbonate	HCO_3^-
hydroxide	OH^-
nitrate	NO_3^-
permanganate	MnO_4^-
phosphate	PO_4^{3-}
sulfate	SO_4^{2-}

Polyatomic Ions

Some ions are made up of several non-metallic atoms joined together. These are called **polyatomic ions** ("poly" means "many"). Consider the hydroxide ion, whose formula is given in Table A2.5. In the compound NaOH, for example, the sodium has a charge of 1+. The oxygen and hydrogen together form the polyatomic hydroxide ion, OH^-, with a charge of 1−. Human bones

contain calcium phosphate. The phosphate anion, PO_4^{3-}, behaves like a single ion with a charge of 3−.

Table A2.5 gives some examples of common polyatomic ions. You can find a more extensive list in Table E in Student Reference 12. Note that in the table, the formula is shown with its ion charge. This is the correct way to show polyatomic ions.

Suffixes for Polyatomic Ions

The two most common suffixes used in naming polyatomic ions are "-ate" and "-ite." When a pair of similar ions exist, such as SO_4^{2-} and SO_3^{2-}, "-ate" and "-ite" are used in their names to differentiate them. SO_4^{2-} is sulfate, and SO_3^{2-} is sulfite. As you may have guessed, "-ate" means more oxygen atoms, and "-ite" means fewer oxygen atoms are part of the ion. But these suffixes do not tell you how many oxygen atoms are actually in the formula. If there are more than two similar ions, then other naming variations are used. Consider for example, the series of chlorine and oxygen ions in Table A2.6. It is not necessary at this stage to memorize all the suffix patterns. Use the ion chart of Table A2.5 and Table E in Student Reference 12 as needed.

Naming Compounds Containing Polyatomic Ions

Naming a compound containing polyatomic ions is simple. Look at the formula, and name the cation, followed by the anion. You do not need to change the ending of a polyatomic ion's name. The only difficulty sometimes is recognizing the ions within the formula. Table A2.7 gives some examples of compounds containing polyatomic ions.

TABLE A2.6 Names of Ions Made up of Chlorine and Oxygen

Ion Name	Ion Formula
perchlorate	ClO_4^-
chlorate	ClO_3^-
chlorite	ClO_2^-
hypochlorite	ClO^-

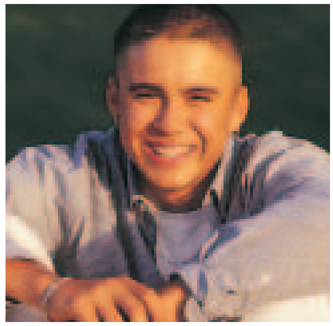

FIGURE A2.18 You probably know that your teeth contain calcium. The calcium is in a calcium phosphate mineral called hydroxyapatite. It includes the polyatomic ion PO_4^{3-}.

TABLE A2.7 Examples of Compounds Containing Polyatomic Ions

Formula	Name	Hints for Writing Names
$Ca(OH)_{2(s)}$	calcium hydroxide	The polyatomic ion is often found inside brackets. Look up OH in Table A2.5 to find its name. It is listed as OH^-, and its name is hydroxide.
$Na_3PO_{4(s)}$	sodium phosphate	Inspect the compound for polyatomic ions. This compound contains the polyatomic ion PO_4^{3-}. Its charge is 3–, and its name is "phosphate."
$(NH_4)_2SO_{4(s)}$	ammonium sulfate	Look in the brackets to find the polyatomic ion NH_4. Look in Table A2.5 to find its name: ammonium. The ammonium ion is the only common polyatomic cation. Memorize the name and formula of ammonium, NH_4^+. The second half is SO_4 which appears in the table as SO_4^{2-}, with the name "sulfate."
$NH_4HCO_{3(s)}$	ammonium hydrogencarbonate	There are no brackets here. Remember that the formula unit for an ionic compound always consists of positive and negative ions in the smallest whole-number ratio that results in a neutral unit. The cation is ammonium, NH_4^+. The second part is the anion, HCO_3^-. Table A2.5 shows HCO_3^-, the hydrogencarbonate ion.

Writing Formulas for Compounds Containing Polyatomic Ions

The rules for writing the formulas for these compounds are similar to the rules for other ionic compounds. One difference is that brackets may be used to show the ratio of the ions. The subscript outside the brackets applies to all the elements inside the brackets. For example, in $Fe_2(SO_4)_{3(s)}$ the subscript 3 with $(SO_4)_3$ means that there are 3 SO_4^{2-} ions for every 2 Fe^{3+} ions in the compound. It also indicates that 3 sulfur atoms and 12 (4 × 3) oxygen atoms are in one formula unit of sulfate.

Example Problem A2.3

What is the formula for iron(III) sulfate?

1. Identify the ions and their charges. Check Table A2.5 for the charge of the polyatomic ion.	iron(III): Fe^{3+} sulfate: SO_4^{2-}
2. Determine the total charges needed to balance.	Fe^{3+}: 3 + 3 = 6 SO_4^{2-}: 2 + 2 + 2 = 6
3. Note the ratio of cations to anions.	2 to 3
4. Use brackets and subscripts to write the formula.	$Fe_2(SO_4)_{3(s)}$

The formula for iron(III) sulfate is $Fe_2(SO_4)_{3(s)}$.

Example Problem A2.4

What is the formula of ammonium dihydrogenphosphate?

1. Identify the ions and their charges. Check Table A2.5 for the charge of the polyatomic ion.	ammonium: NH_4^+ dihydrogenphosphate: $H_2PO_4^-$
2. Determine the total charges needed to balance.	NH_4^+: 1 $H_2PO_4^-$: 1
3. Note the ratio of cations to anions.	1 to 1
4. Write the formula. Brackets and a subscript are not needed when there is only one ion.	$NH_4H_2PO_{4(s)}$

The formula of ammonium dihydrogenphosphate is $NH_4H_2PO_{4(s)}$.

Practice Problems

Use Table A2.5 to help you.

3. Write the formulas of the following ionic compounds:
   a) barium hydroxide
   b) iron(III) carbonate
   c) copper(I) permanganate

4. Write the names of the following ionic compounds:
   a) $Au(NO_3)_{3(s)}$
   b) $(NH_4)_3PO_{4(s)}$
   c) $K_2Cr_2O_{7(s)}$

Molecular Compounds

The burning methane hydrate ice described in the Exploring section at the beginning of this unit is made up of two important compounds: methane ($CH_{4(s)}$) and water ($H_2O_{(s)}$). Methane is the chemical name for natural gas. The electricity that operates your computer, CD player, and lights may be generated with methane. Your home may be heated with methane, and your dinner may be cooked with it. The other compound, water, is essential to life. It covers about 74% of Earth. You drink it, wash in it, and play in it. It's found in every living thing, from one-celled organisms and plants to insects and mammals.

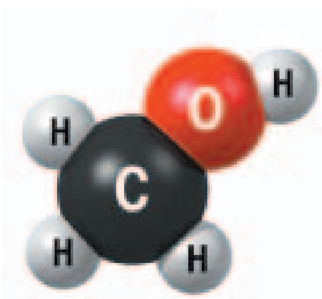

FIGURE A2.19 A methanol molecule. Methanol is the antifreeze used in car and truck engines.

Neither of these compounds contains any ions—they are each made up of molecules. A **molecule** forms when two or more non-metallic atoms bond together. It can be made up of atoms of different elements or of atoms of all the same element. For example, $CH_3OH_{(l)}$ (methanol) and $O_{2(g)}$ (oxygen gas) are both molecules. Figure A2.19 shows the structure of a methanol molecule.

Recall that the formula unit of an ionic compound represents a ratio of ions in a crystal lattice. In a solid ionic compound, this lattice extends in all directions. A formula unit is not an independent unit—it is just one part of a crystal lattice. Molecules are independent units made up of fixed numbers of atoms bonded together.

Unlike ionic compounds, molecular subtances can be solid, liquid, or gas at room temperature. They tend to be poor conductors of electricity, even in solution. Many do not dissolve in water very well. You will learn more about the properties of molecular compounds in section A2.3. Table A2.8 describes some common compounds.

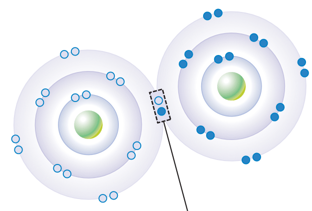

FIGURE A2.20 A molecule of chlorine gas forms when two chlorine atoms share one pair of electrons to form a covalent bond.

TABLE A2.8 Examples of Common Molecular Compounds

Common Name	Formula	Chemical Name	Application
sugar	$C_{12}H_{22}O_{11(s)}$	sucrose	sweetener
alcohol	$CH_3CH_2OH_{(l)}$	ethanol	component of alcoholic beverages
nail polish remover	$CH_3COCH_{3(l)}$	acetone	solvent
natural gas	$CH_{4(g)}$	methane	heating fuel

Sharing Electrons—Covalent Bonds

The atoms in a molecule are joined together by **covalent bonds** that form when atoms share electrons. Each pair of shared electrons forms one covalent bond. Electrons are not transferred from one atom to another as they are in ionic bonds.

Chlorine gas is an example of a substance that has molecules formed of only one element. Each chlorine molecule is made up of two chlorine atoms joined by a covalent bond. Recall that a chlorine atom has seven valence electrons in its outer energy level. For this outer energy level to be filled, an additional electron is needed. A molecule of chlorine gas is created when two atoms of chlorine each share an electron to form a covalent bond. Figure A2.20 shows how sharing an electron forms a covalent bond.

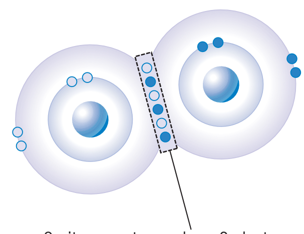

FIGURE A2.21 A molecule of nitrogen gas forms when two nitrogen atoms share three pairs of electrons to form covalent bonds. Each atom now has a stable outer energy level of eight electrons.

Notice that the electron is not transferred from one atom to another as it would be in an ionic bond. In covalent bonds, atoms share electrons so that their outer energy levels become filled. Other examples of covalent compounds include $H_2O_{(l)}$ (water), $NH_{3(g)}$ (ammonia), and $C_{12}H_{22}O_{11(s)}$ (sugar).

In some covalent compounds, the atoms share more than two electrons. For example, nitrogen gas occurs in the form $N_{2(g)}$. (Nitrogen gas makes up 78% of our atmosphere.) An atom of nitrogen has five valence electrons. To form $N_{2(g)}$, two nitrogen atoms share three pairs of electrons (Figure A2.21). In carbon dioxide ($CO_{2(g)}$), all atoms share two pairs of electrons (Figure A2.22).

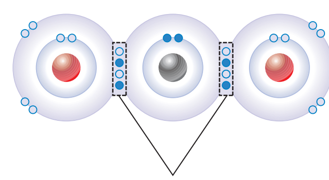

FIGURE A2.22 A molecule of carbon dioxide gas forms when one carbon atom and two oxygen atoms each share two pairs of electrons to form covalent bonds.

reSEARCH

Mercury is unusual because it is a metallic liquid at room temperature. The mercury(I) ion is also unusual in that it exists as a polyatomic ion. Find out the formula of this ion and some of its properties. You can extend this research further by finding out about calomel electrodes. Begin your search at www.pearsoned.ca/school/science10

Molecular Elements

Nitrogen is known as a **molecular element** because it forms molecules made up of only one type of atom. Its molecules are **diatomic**, which means each one is composed of only two atoms: $N_{2(g)}$ ("di-" means "two"). Some elements form polyatomic molecules. For example, sulfur forms a ring of eight atoms and has the formula $S_{8(s)}$. Other elements are monatomic. Their atoms can exist on their own. Carbon, for example, is written as $C_{(s)}$, although in both diamond and graphite it is connected to other carbon atoms in very large arrays. Formulas of molecular elements are summarized in Table A2.9.

TABLE A2.9 The Chemical Formulas of Molecular Elements

Monatomic	$C_{(s)}$	noble gases		all metals			
Diatomic	$H_{2(g)}$	$N_{2(g)}$	$O_{2(g)}$	$F_{2(g)}$	$Cl_{2(g)}$	$Br_{2(l)}$	$I_{2(s)}$
Polyatomic	$O_{3(g)}$(ozone)	$P_{4(s)}$	$S_{8(s)}$				

It can be very helpful to have the diatomic elements memorized, particularly when writing chemical reactions, as you will do in the next section. One way to remember them is that the diatomic elements are the "gens": hydro*gen*, nitro*gen*, oxy*gen*, and the halo*gens*.

Molecular Compounds That Do Not Contain Hydrogen

A binary compound contains two elements. Some of these compounds contain hydrogen, and some do not. IUPAC rules for naming binary molecular compounds not containing hydrogen are similar to the rules for naming ionic compounds. For molecules, Greek prefixes are used to indicate how many atoms of each element are present in the compound. For example, $P_4O_{10(s)}$ is called tetraphosphorus decaoxide: "tetra" means "4," and "deca-" means "10." Table A2.10 lists the prefixes used for naming binary compounds.

TABLE A2.10 Prefixes Used for Naming Binary Molecular Compounds

Prefix	Number
mono	1
di	2
tri	3
tetra	4
penta	5
hexa	6
hepta	7
octa	8
ennea (nona)	9
deca	10

Any compound that does not have a metal or an ammonium ion in its formula is molecular. The format for naming binary molecular compounds not containing hydrogen is:

prefix + first element *followed by* prefix + second element ending in "-ide"

Note that the prefix "mono-" is not used when the first element is only one atom. When the prefix "mono-" is required before "oxide," the last "o" in the prefix is usually dropped. For example, it is "monoxide," not "monooxide." Table 2.11 shows how the format is used in two examples.

TABLE A2.11 Examples for Naming Binary Molecular Compounds

Steps	Examples	
	$N_2O_{(g)}$	$PBr_{3(g)}$
1. Name the first element.	nitrogen	phosphorus
2. Name the second element with "-ide" at the end.	oxide	bromide
3. Add prefixes indicating numbers of atoms.	dinitrogen monoxide	phosphorus tribromide

Here are some examples:

$CO_{(g)}$	carbon monoxide
$SO_{2(g)}$	sulfur dioxide
$CS_{2(g)}$	carbon disulfide
$N_2O_{3(g)}$	dinitrogen trioxide
$CCl_{4(l)}$	carbon tetrachloride
$P_4O_{10(s)}$	tetraphosphorus decaoxide

Example Problem A2.5

Write the name of the compound that has the formula $PCl_{5(s)}$.

The rules listed above are applied in order. The first element is phosphorus. The second element is chlorine, so this compound is a chloride. Since only one atom of phosphorus is present, no prefix is used. Five atoms of chlorine are present, so the prefix is "penta."

The name of the compound with the formula $PCl_{5(g)}$ is phosphorus pentachloride.

Practice Problem

5. Write the names or formulas for the following molecular compounds:

a) $CO_{2(g)}$
b) $N_2O_{(g)}$
c) $PCl_{3(g)}$
d) oxygen difluoride
e) dinitrogen tetrasulfide
f) sulfur trioxide

Molecular Compounds That Contain Hydrogen

Hydrogen is unique in many ways, and this is reflected in naming systems. Many compounds containing hydrogen have simply been given names. The name "water," for example, was chosen by IUPAC to be the official name for H_2O. These names have to be memorized or found by referring to a chart like Table A2.12. Note, in particular, that the prefix "mono-" is omitted in $H_2S_{(g)}$, which is named hydrogen sulfide.

Writing the formulas for molecular compounds is easy because the prefixes in the names indicate the number of each element. However, predicting formulas when elements combine is difficult because more than one combination is possible; for example, $CO_{(g)}$ or $CO_{2(g)}$.

TABLE A2.12 Examples of Names of Molecular Compounds Containing Hydrogen

IUPAC Name	Formula and State at 25°C
water	$H_2O_{(l)}$
hydrogen peroxide	$H_2O_{2(l)}$
ammonia	$NH_{3(g)}$
sucrose	$C_{12}H_{22}O_{11(s)}$
methane	$CH_{4(g)}$
propane	$C_3H_{8(g)}$
methanol	$CH_3OH_{(l)}$
ethanol	$C_2H_5OH_{(l)}$
hydrogen sulfide	$H_2S_{(g)}$